RESEARCH ARTICLE

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Investigating the Enthalpy of Dissolution of Ionic and Polar Substances in Water

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Abstract

Since water is a polar molecule with a bent structure, it dissolves many ionic and polar substances. Ionic solids are mostly soluble in water and water is the most common solvent of known ionic compounds. In an ionic solid, ions with opposite charges are held together through electrostatic attraction. The solubility of a substance in a solvent at a certain temperature is the maximum amount of that substance that dissolves in a certain volume of that solvent. As a result of the dissolution of one substance in another substance, three steps occur: separation of dissolved particles, separation of solvent particles, and mixing of solute particles and solvent particles. The first and second stages are endothermic due to the separation of the molecules of the solute substance and the solvent from each other, and the third stage is exothermic due to the merging of the molecules of the solute substance and the solvent. The enthalpy of dissolution of one mole of solute in a large amount of solvent is called the enthalpy of dissolution of the same substance. The dissolution process is associated with energy exchange, which is obtained by using the enthalpies of the solute substance, solvent, and hydration enthalpy in the form of enthalpy of solution. The enthalpy of the solution is also determined using the Brunhaber cycle. If the set of mentioned steps gives numbers with a positive sign, the process is endothermic and if the numbers with a negative sign are obtained, the process is exothermic. Scientists use this property of dissolution of materials to produce heat-generating and cold-generating packages, which are used for different purposes. The methods used in collecting information to achieve the desired goal are of the library type and quantitative and qualitative methods are also used.

Keywords: water, solubility, solid, liquid, and ice.

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Introduction

Matter is generally found in nature in three forms: gas, liquid and solid. A state of matter whose constituent particles are separated from each other and irregular is called gas, particles that are close together but irregular, liquid and particles that are close to each other and in an orderly manner are called solid. In crystalline solids, ions are placed in certain places of a crystal lattice and their movements are limited by vibration around fixed points. The kinetic energy of the molecules in the crystal is very weak, so the attraction forces act strongly and create regular crystal structures. Regular crystal structures, whose constituent particles are formed in three dimensions, are called the crystal lattice of that object. The formation of the crystal network is associated with the release of energy, and the energy of the network is the amount of energy released during the formation of an ionic solid mole from its constituent gas ions. The enthalpy of formation of a solid network is always positive and its energy can be measured using the Haber-Bron cycle. When a crystalline substance enters a liquid such as water, it begins to dissolve, resulting in a dissolution process and the corresponding solution is produced. Solutions are homogenous compounds that result from the fusion of dissolved and solute substances in which the substance's origin has not changed and its physical and chemical properties are the same in all its portions. The dissolution of one substance in another substance is done in three stages, which are related to the separation of the dissolved substance, the solvent, and the mixing of the dissolved substance and the solvent. Solubility is a fundamental concept in the science of chemistry and the topic of solutions, which expresses the degree of dissolution of a substance in a specific solvent. The enthalpy related to

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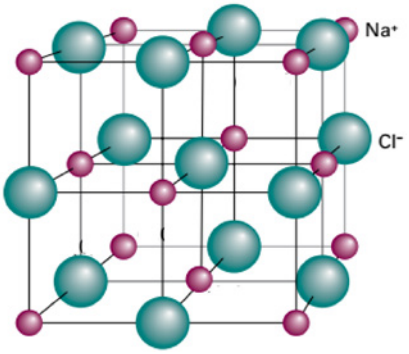

Figure 1: Representation of ions in sodium chloride crystal (Atkins, 2018, p. 650).

the dissolution of one mole of solute in a large amount of solvent is called the enthalpy of dissolution, which is obtained from the integration of the enthalpy change of three steps. If the enthalpy of the mixture and the enthalpy of the solute are greater than the enthalpy of the solution, the process is exothermic and involves the release of heat. But if the mixed enthalpy is smaller than the soluble enthalpy, the process is endothermic and is associated with energy absorption, and these absorbed or released energies in the process of dissolving substances in a solvent are shown as enthalpy of solution. Today, by using the enthalpy of dissolution of materials, they produce heat packs and cooler packages, which are of particular importance in various especially medicine and sports.

The purpose of this research is to know the solubility and enthalpy of dissolution of substances, and the understanding of these properties is considered an urgent necessity for students and teachers, the reason why some substances release heat and some absorb heat when dissolved. It is included, it should be turned on. And on the other hand, the reason for sprinkling salt and soil when freezing the roads to melt the ice and using small grains of salt in the ice cream makers to close the ice cream in the ice cream maker, which is a dissolution process, should be clarified.

The methods used in this research are quantitative and qualitative, and in the form of a review of a library of reliable scientific books, they have been used to enrich as much as possible.

Ionic solids

The kinetic energy of gas molecules decreases as the temperature decreases. As a result, if the gas is cooled enough, the intermolecular forces of attraction cause the molecules to condense and turn into a liquid. In the liquid, the molecules are closer to each other and the force of attraction between them is greater than in the gas state. In this way, the molecular movement in the liquid state is more limited than in the gas state. Further cooling reduces the kinetic energy of molecules and leads to their freezing. In a crystalline solid, the molecules are placed in places of a crystal lattice and their movements are limited to oscillation around these fixed points. The kinetic energy of the molecules in the crystal is so weak that the forces of attraction easily prevail and create very regular crystal structures (Mortimer, 2010, p. 158).

The arrangement of ions in ionic compounds is a repeating pattern, and each ion in its place interacts with several ions that have the opposite charge. The arrangement of ions in a salt crystal follows a specific pattern depending on the relative sizes of the cation and anion, and this pattern is repeated throughout the crystal. The structure that is formed by arranging the constituent particles of an object in three dimensions is called the crystal lattice of that object.

The formation of the crystal network is associated with the release of energy. Lattice energy is the amount of energy released during the formation of an ionic solid mole from its constituent gaseous ions. In the case of sodium chloride, its lattice energy is $+787.5\,\mathrm{kJ\,mol^{-1}}$ (Abedini et al., 2008, p. 59).

$$\mathrm{Na^{+}(g) + Cl^{-}(g) \rightarrow NaCl(s) \quad +787.5\,kJ\,mol^{-1}}$$

The enthalpy of formation of a solid can be divided into different components and determined using the Haber-Bron cycle. To determine the enthalpy of the NaCl network, its general interaction can be considered as the final result of five steps as follows:

$$\mathrm{Na(s)+\tfrac{1}{2}Cl_2(g) \rightarrow NaCl(s)}$$

$$\mathrm{Na(s) \rightarrow Na(g)} \qquad \Delta \mathrm{H}^{\phi}_{\mathrm{sub}}\mathrm{(Na)}.1$$

$$\mathrm{Na(g) \rightarrow Na^{+}(g)+e^{-}(g)} \qquad \Delta \mathrm{H}^{\phi}_{\mathrm{i}}\mathrm{(Na)}.2$$

$$\mathrm{\tfrac{1}{2}Cl_2(g) \rightarrow Cl(g)} \qquad \tfrac{1}{2}\Delta \mathrm{H}^{\phi}\mathrm{(Cl\text{-}Cl)}.3$$

$$\mathrm{Cl(g)+e^{-}(g) \rightarrow Cl^{-}(g)} \qquad \Delta \mathrm{H}^{\phi}_{\mathrm{ea}}\mathrm{(Cl)}.4$$

$$\mathrm{Na^{+}(s)+Cl^{-}(g) \rightarrow NaCl_2(s)} \qquad .5$$

The enthalpy change dependent on step 5 is the negative lattice enthalpy $(\Delta \mathrm{H}_L)$. Lattice enthalpy in the general case is the standard enthalpy change related to the formation of a gas containing ions from the desired crystalline solid:

$$\mathrm{MX(s) \rightarrow M^{+}(g) + X^{-}(g)} \;\; \Delta \mathrm{H}^{\phi}_{L}$$

All network enthalpies are positive according to Table 1.

Table 1:network enthalpies $\ddot{\mathrm{A}}\mathrm{H}^{\phi}_{L}\,/(\mathrm{kJ\,mol^{-1}})$ at 298K .

NaF	926	MgS	3406
NaCl	787	$MgCl_2$	2524
KCl	717	$CaCl_2$	2255
MgO	3850	CaO	3461

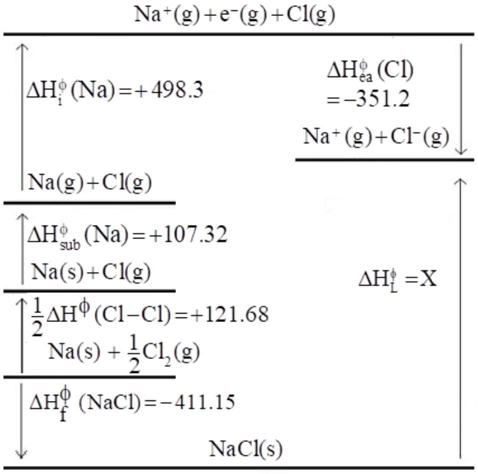

Figure 2: Born-Haber cycle to determine NaCl(s) lattice enthalpy.

Lattice energy is the same as lattice enthalpy at T = 0. The desired sequential steps are shown in Figure 2.

The importance of a cycle is that the sum of the enthalpy changes of the participants is zero (because the initial and final states are the same). From that, we understand that if all the enthalpy changes of the cycle participants except one are known, the unknown enthalpy can be determined.

Example: Calculate the lattice enthalpy of $\mathrm{NaCl(s)}$ in $298\mathrm{K}$ using Boron-Haber cycle in (Figure 2)?

Solution: x should be obtained using the condition that in figure 2 the distance from the bottom to the top on the left or right side is equal, it can be obtained with the data obtained below.

$\frac{1}{2}\text{ÄH}^{\phi}(\mathrm{Cl\text{-}Cl}) = +121.6\,\mathrm{kJ\,mol^{-1}}$

$\text{ÄH}^{\phi}_{\mathrm{sub}}(\mathrm{Na,s}) = 107.32\,\mathrm{kJ\,mol^{-1}}$

$\text{ÄH}^{\phi}_{\mathrm{i}}(\mathrm{Na}) = +498.3\,\mathrm{kJ\,mol^{-1}}$

$\text{ÄH}^{\phi}_{\mathrm{ea}}(\mathrm{Cl}) = -351.2\,\mathrm{kJ\,mol^{-1}}$

$\text{ÄH}^{\phi}_{f}(\mathrm{NaCl_2,s}) = -411.15\,\mathrm{kJ\,mol^{-1}}$

On the left side, the distance from $\mathrm{NaCl_2(s)}$ to the top of the cycle is equal to:

$$(411.15+121.68+107.32+498.3)\,\mathrm{kJ\,mol^{-1}} = 1138.4\,\mathrm{kJ\,mol^{-1}}$$

On the right side, the distance from the bottom to the top is $(X+351.2)\,\mathrm{kJ\,mol^{-1}}$. Two distances are equal, so $X=782.2$. Because the lattice enthalpy is positive, it is $\text{ÄH}^{\phi}_{L}=787.2\,\mathrm{kJ\,mol^{-1}}$ (Atkins, 2015, p. 64).

Solubility of ionic solids

Since water is a polar molecule with a curved structure, it dissolves many ionic and polar substances in itself and a number of other molecules that are not ionic and polar, but have a hydrogen bond, dipole-dipole induced forces and they are the force of dispersion, it also dissolves in itself (Silberberg, 2011, p. 5).

Partial frames of molecules act as small electric fields that create a specific orientation in a liquid or solid (Jalili, 2007, p. 75).

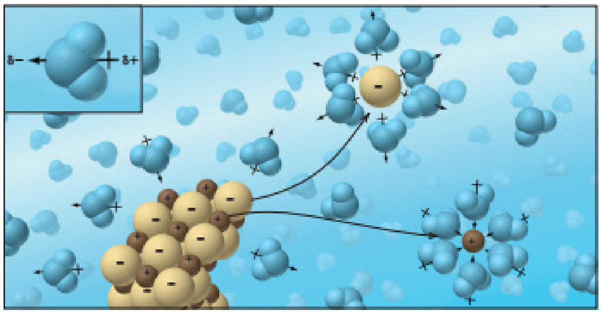

Figure 3: Dissolving sodium chloride in water

In an ionic solid, ions with opposite charges are held together by electrostatic attraction. Water separates ions by replacing this attraction with the attraction between water molecules and ions. Imagine a grain of an ionic compound surrounded by bent polar water molecules. The negative end of some water molecules is attracted to the cations and the positive end of the other is attracted to the anions (Silberberg, 2011, p. 145).

The attraction that occurs between bipolar molecules and ions in this way causes the ions to separate from the crystal and float in the liquid phase. Solved ions are hydrated and surrounded by a coating of water molecules and move in the solution (Mortimer, 2010, p. 181).

Although most ionic compounds are soluble in water, many others are insoluble. In the second case, the electrostatic attraction between compound ions remains more than the attraction between ions and water molecules, so the solid remains intact to a large extent (Jalili, 2007, p. 76).

Usually, the component of a solution that is greater than the other components is called the solvent and the other components are called the solute. Some substances dissolve in each other in any proportion. The complete miscibility of the components of all gaseous solutions and some solid and liquid solutions is one of their important features (Mortimer, 2010, p. 180).

As you know, there is a limit for the dissolution of various substances in a certain amount of water at room temperature. This limitation is determined by the solubility or ability to dissolve that substance (Shoray talif groh shemy, 2008, p. 22).

The solubility of a substance in a specific solvent and at a certain temperature is the maximum amount of that substance that dissolves in a certain amount (100g of water) of that solvent and creates a stable system (Mortimer, 2010, p. 180).

Water is the most abundant and common known solvent that dissolves many ionic and covalent compounds. Solutions whose solvent is water are called aqueous solutions. Water and aqueous solutions play an important role in daily life and most vital processes such as digestion, absorption and metabolism of food are carried out in aqueous solution. In chemical industries, a large number of interactions take place in the water environment. Although water is the most important known solvent, there are other important organic

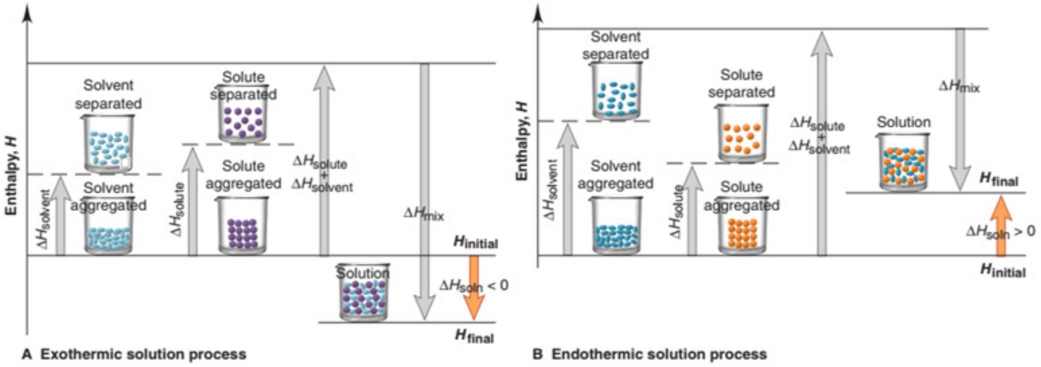

Figure 4: The solution formation cycle and the enthalpy of the components and the heat of the solution.

solvents such as hexane, ethanol, and acetone. The solutions obtained from organic solvents are called non-aqueous solutions (Ruh Elahi et al., 2009, p. 77).

Imagine a solvent and solutes that want to form a solution. Both contain particles that attract each other. For a substance to dissolve in another substance, three things must happen: (1) the solute particles must be separated from each other, (2) some of the solvent particles must be separated to create space for the solute particles, (3) Solvent and solute particles mix together. To separate the particles regardless of the attractions in the solute and the attractions in the solvent, some energy is absorbed and some energy is released as a result of mixing and absorbing the molecules with each other. As a result of these changes, the solution formation process is usually accompanied by an entropy change. This process can be divided into three stages:

- Separating soluble particles from each other. Because intermolecular attractions must be

overcome in this stage, this stage is endothermic:

Solute (aggregated) + heat → solute (separated) $\Delta H_{solute} > 0$

- Separating the solvent particles from each other. At this stage, the attractions must be overcome, so this stage is also endothermic.

Solvent (aggregated) + heat → solvent (separated) $\Delta H_{solvent} > 0$

- Soluble and solvent particles are mixed. The particles attract each other, so this step is exothermic:

Solute (separated) + solvent (separated) → solution + heat $\Delta H_{mix} < 0$

The enthalpy change of solution formation from solute and solvent is the enthalpy of solution (ΔH_{soln}) and it is obtained by integrating the enthalpy change of three steps. This entire process is called a solution thermochemical cycle, which is another application of Hess's law:

$$\Delta H_{soln} = \Delta H_{solute} + \Delta H_{solvent} + \Delta H_{mix}$$

If the sum of the terms $\Delta H_{solvent}$ and ΔH_{solute} is smaller than the term ΔH_{mix}, it is negative and the process (ΔH_{soln}) will be exothermic. (Figure 4A) shows the enthalpy graph for the formation of such a solution. If the sum of the terms $\Delta H_{solvent}$ and ΔH_{solute} including ΔH_{mix} is greater, ΔH_{soln} will be positive and the process will be endothermic (Figure 4B). If ΔH_{soln} is positive and large, it is possible that the solute is not significantly dissolved in the solvent (Silberberg, 2013, p. 397).

Enthalpy of formation of objects in solution

Thermodynamic data are often reported in terms of the enthalpy of a compound relative to its constituent elements at standard conditions. Therefore, the standard enthalpy of formation (ΔH_f^{ϕ}) of a body is the standard enthalpy of interaction in which that body is formed from its constituent elements that are in their reference states. The reference state of an element is its most stable state at a certain temperature and pressure of 1 bar (Atkins, 2015, p. 63).

The enthalpy of formation of a substance in solution is the standard enthalpy change that depends on the following interactions:

$$\tfrac{1}{2}H_2(g) + \tfrac{1}{2}Cl_2(g) \rightarrow HCl(aq) \quad \Delta H_f^{\phi}(HCl,aq) = -167\,kJ\,mol^{-1}$$

$$Na(s) + \tfrac{1}{2}Cl_2(g) \rightarrow NaCl(aq) \quad \Delta H_f^{\phi}(NaCl,aq) = -407\,kJ\,mol^{-}$$

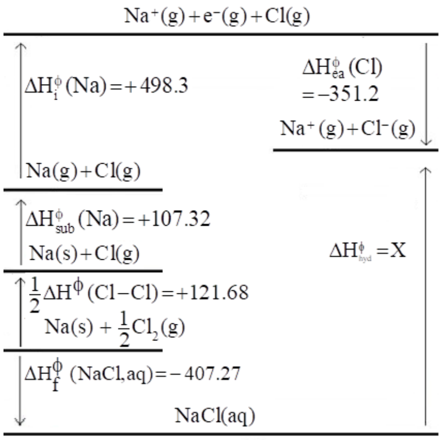

Figure 5: The same thermodynamic cycle for determining the enthalpy of hydration of Na^+ and Cl^- ions.

Table 2: Standard molar enthalpies of hydration at infinite dilution $\ddot{A}H^{\phi}_{hydr}$ /(kJ mol^{-1}) at 298K

	Li^+	Na^+	K^+
F^-	-1026	-911	-828
Cl^-	-884	-783	-685
Br^-	-856	-742	-658

By using the Bron-Haber cycle for solids, it is possible to analyze the enthalpy of formation in solution. The only difference is that instead of the network formation step (step 5 in Figure 2) of

$$5.\ Na^+(g) + Cl^-(g) \rightarrow Na^+(aq) + Cl^-(aq) = NaCl(aq)$$

Used. In this case, the enthalpy change is as the enthalpy of hydration (ΔH^{ϕ}_{hydr}) of gaseous ions (in the general case, the enthalpy of solvent coverage is (ΔH^{ϕ}_{solv})). Since all the other enthalpies participating in the cycle are known, the mentioned value can be determined: the distance from the bottom to the top on the left side of Figure 5 $(1134.6 kJ\,mol^{-1})$ is equal to the similar distance on the right side $(351.2 kJ\,mol^{-1} + X)$, so, $X = 783.4 kJ\,mol^{-1}$ is a solvent process. The hydration is a downward step on the right, so the enthalpy of hydration of NaCl is $-x$ or $-783.4 kJ\,mol^{-1}$.

Other values can also be determined in the same way, some of which are collected in Table 1. These values are related to the formation of a solution at zero concentration (infinite dilution) because the reactions between ions that are important in normal concentrations are not taken into account here (Atkins, 2015, p. 67).

Enthalpy of formation of individual ions in solution

The enthalpy of formation of a completely ionized compound in solution can be considered as the sum of the enthalpies of formation of the solutions of the constituent ions of that compound. Thus, the enthalpy of $HCl(aq)$ formation can be considered as the sum of the enthalpies of $H^+(aq)$ and $Cl^-(aq)$ formation. Now the problem is to decide how to divide the total amount between the two types of ions. By convention, the standard enthalpy of formation of $H^+(aq)$ has been chosen to be zero.

At all temperatures $\frac{1}{2}H_2(g) \rightarrow H^+(aq) \quad \Delta H^{\phi}_f\ (H^+,aq)$

From the above agreement, the enthalpy of formation of $Cl^-(aq)$ is $-167 kJ\,mol^{-1}$. By combining this value with the enthalpy obtained for the formation of $NaCl(aq)$, we can obtain a value for the formation of $Na^+(aq)$ and continue the work in the same way. The values collected in Table 2 were obtained in the same way.

Table 3: Limiting enthalpies of formation of ions in $\ddot{A}H^{\phi}_f$ /(kJ mol^{-1}) at 298K aqueous solution.

Cations		Anions	
H^+	0	OH^-	-230.0
Na^+	-240.1	Cl^-	-167.2
Cu^{2+}	+64.8	SO_4^{2-}	-909.3
Al^{3+}	-531	PO_4^{3-}	-1277.4

As we analyzed the enthalpy of formation of NaCl in terms of a thermodynamic cycle, we can repeat the same work for the enthalpy of hydration of individual ions. However, here we can no longer accept an optional convention about the proton hydration enthalpy, because it would not be obtained with our previous choice the enthalpy of formation. We set the value of zero for the enthalpy regarding of the $\frac{1}{2}H_2(g) \rightarrow H^+(aq)$ step, we cannot choose zero for $H^+(g) \rightarrow H^+(aq)$ at the same time. Therefore, we have to estimate the actual amount of enthalpy change in this step by calculating the interaction energy of a proton with the surrounding water molecules. Several consensuses (from mass spectrometry and proton binding energy to a small group of water molecules) about $H^+(g) \rightarrow H^+(aq)\ \Delta H^{\phi} \approx -1090 kJ\,mol^{-1}$ are at work. By accepting this value, the desired value for $Cl^-(g) \rightarrow Cl^-(aq)$ can be obtained from it and the related data of HCl. Then, by combining the value of Cl^- and $\ddot{A}H_{hyd}$ obtained for NaCl, a value of about $-400 kJ\,mol^{-1}$ was concluded for $Na^+(g) \rightarrow Na^+(aq)$. The hydration enthalpy of the single ions collected in Table 3 were obtained in the same way.

The data in Table 3 are generally confirmed. Hence, the small, high-charge ions have the most negative hydration enthalpies. Because they strongly absorb solvent molecules. Electron hydration enthalpy means process enthalpy $e^-(g) \rightarrow e^-(aq)$ is only about $-170 kJ\,mol^{-1}$. From this value, it can be concluded that when an electron is trapped in water (such as when water is affected by high-energy radiation), the electron occupies a volume of about two water molecules.

Calculation of enthalpy of dissolution of ionic solids

The process of dissolution is associated with energy exchange, and it is determined using enthalpy of dissolution.

Table 4: Enthalpy of standard hydration of ions $\ddot{A}H^{\phi}_{hydr}$/(kJ mol^{-1}) at 298K

Li^+	-520	F^-	-506
Na^+	-405	Cl^-	-364
K^+	-321	Br^-	-337

Enthalpy change related to the dissolution of one mole of solute in a large amount of solvent; it is called heat of dissolution or enthalpy of dissolution ($\Delta H_{dissoln}$) (Rouh Elahi et al., 2009, p. 81).

It is difficult to measure the enthalpy of solution components ($\Delta H_{solvent}$ and ΔH_{mix}) separately. The combination of these two forms the enthalpy change in the mixing process, solvent coating, which is the process of surrounding the solute with solvent particles. Solvation in water is called hydration. Therefore, the enthalpy of hydration is obtained from the enthalpy changes for separating water molecules ($\Delta H_{solvent}$) and mixing the solute with them (ΔH_{mix}).

$$\Delta H_{soln} = \Delta H_{solute} + \Delta H_{\text{hydr of the ions}}$$

The heat of hydration is the key factor for ionic solid dissolution. The formation of strong ion-dipole forces creates heat that is more than what is needed to break hydrogen bonds, so hydration is an exothermic ion. ΔH_{hydr} of an ion is defined as the enthalpy change of hydration for 1mol of ion (gas):

$$M^{+}(g) \text{ or } X^{-}(g) \longrightarrow M^{+}(aq) \quad \text{or} \quad X^{-} \quad (aq) \qquad \Delta H_{\text{hydr of the ions}}(\text{always} > 0)$$

The heat of hydration has a trend based on the density of the charge ion and, the ratio of the charge ion to its volume. In general, the higher the impurity of the charge ion, the larger and negative $\ddot{A}H_{hydr}$ will be. According to Coulomb's law, the larger the charge of the ion and the closer it is to the opposite pole of the water molecule, the stronger the attraction. Therefore, an ion with a 2+ charge (such as Ca^{2+}) attracts water molecules more strongly than an ion with a 1+ charge (such as Na^{1+}).

A smaller 1+ ion (such as Li^{1+}) attracts water molecules more strongly than a larger ion (such as Cs^{1+}).

The energy required to separate the ionic solute (ΔH_{solute}) into gaseous ions is the lattice energy ($\Delta H_{lattice}$) of the solute, which is positive and very large.

$$M^{+} X^{-}(s) \rightarrow X^{+}(g) \quad \text{or} \quad X^{-}(g) \qquad \Delta H_{solute} = (\text{always} < 0) \qquad \Delta H_{lattice}$$

Therefore, the heat of solution for ionic citrus in water is obtained from the combination of the lattice energy (always positive) and the heat of hydration of cation and anion (always negative):

$$\Delta H_{soln} = \Delta H_{lattice} + \Delta H_{\text{hydr of the ions}}$$

The sign of heat of solution is determined by the magnitude of these two contributions (Selberberg, 2011, p. 10).

The dissolution of an ionic or covalent compound in water from a molecular point of view consists of three stages. Now we will discuss the dissolution process of a number of ionic compounds including sodium chloride, sodium hydroxide and ammonium nitrite:

- Collapse of the NaCl crystal network: this step is associated with energy consumption, so it is considered exothermic.

  $$NaCl\ (s) \rightarrow Na^{+}(g) + Cl^{-}(g)\ \Delta H = 786 kJ.mol^{-1}$$
- Separation of water molecules from each other: this step is also exothermic.
- Establishing a strong attraction between ions (Na^{+}, Cl^{-}) and water molecules: this stage is endothermic and a lot of energy is released during it.

The sum of steps 2 and 3 is called the hydration step. At this stage, Na^{+} and Cl^{-} ions are surrounded by water molecules. The hydration step is generally exothermic. The enthalpy of dissolution of NaCl in water is the algebraic sum of the heat required for the collapse of the NaCl crystal lattice and the heat released due to the hydration of Na^{+} and Cl^{-} ions.

Example: if the energy required for the collapse of the NaCl crystal lattice (lattice enthalpy, $\Delta H_{lattice}$) is 786 kJ.mol^{-1} and the total energy released in the hydration (enthalpy of hydration, ΔH_{hydr}) of Na^{+} and Cl^{-} ions is -783 kJ.mol^{-1} Calculate the enthalpy of dissolution of NaCl in water (Rouh Elahi et al., 2009, p. 83).

$$\Delta H_{lattice} = NaCl = +786 \text{ kJ mol}^{-1}$$
$$\Delta H_{hydr} = Na^{+} + Cl^{-} = -783 \text{ kJ mol}^{-1}$$
$$\Delta H_{dissoln} = \Delta H \text{ lattice} + \Delta H_{hydr}$$
$$\text{kJ mol}^{-1} = +3 \ )\ (-783 \text{ kJ mol}^{-1} + +786 \text{ kJ mol}^{-1}\ \Delta H_{dissoln} =$$

The enthalpy of NaCl solution is positive and small. Its lattice energy is only slightly larger than the enthalpies of its ions. Therefore, due to the dissolution of NaCl in a water thermos, we do not notice the temperature change (Selberberg, 2011, p. 12).

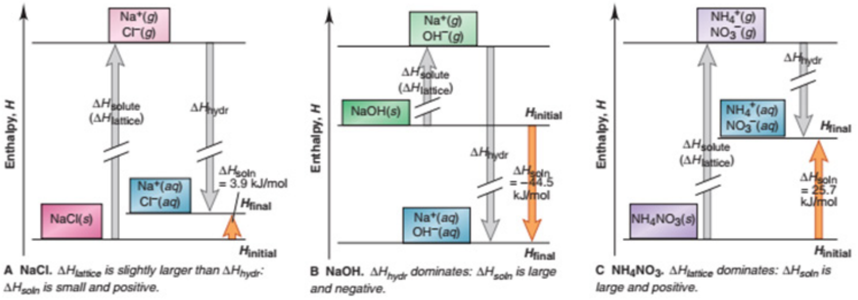

Figure 6: Dissolution of ionic compounds in water. The enthalpy graph of an ionic compound includes $\Delta H_{lattice}$ ($\Delta H_{lattice}$ is always positive) and the sum of heats of hydration of ions (ΔH_{hydr} is always negative) (Silberberg, 2013, p. 399)

Due to the cheapness of salt compared to many other products, the use of this substance in cleaning roads covered with snow in the winter season is increasing day by day. The property of using salt is that salt lowers the freezing point of water. For example, a 23.3% concentration of water and salt solution causes water to freeze at -60°F and a 29.8% concentration of calcium chlorite solution at -65°F. This nature causes both salt and calcium chlorite to be used in winter to melt snow and ice (Soleimani Kermani, 2008, p. 187).

Eutectic, which consists of 23% sodium chloride and 77% water, melts at -21.1^0C. When salt is added to ice under isotherm conditions (for example, when salt is sprinkled on icy roads), the mixture will melt if the temperature is above -21.1^0C. Adding salt to ice under adiabatic conditions (for example, when salt is added to ice in a vacuum thermos) melts the ice, but in doing so it absorbs heat from the rest of the mixture and the temperature of the system decreases. . If enough salt is added, the cooling will continue below the temperature of -21.1^0C (Atkins, 2004, p. 319). But if we dissolve sodium hydroxide (NaOH) in water, I can feel the heating of water thermoses. The lattice energy of NaOH is much smaller than the hydration enthalpy of ions, so the dissolution of NaOH is very exothermic ($\Delta H_{dissoln} = -44.5$ kJ.mol^{-1}). At the end, if I dissolve ammonium nitrite (NH_4NO_3) in water, we can feel the cooling of the water thermos. In this case, the lattice energy is much larger than the enthalpy of ion hydration, so the process is completely endothermic ($\Delta H_{dissoln} = 25.7$ kJ.mol^{-1}) (Silberberg, 2013, p. 399).

Applications of enthalpy of dissolution of solids

One of the ways of exchanging heat with the environment is using exothermic and endothermic chemical or physical interactions. Of course, these interactions must exchange heat in such a way that it is closed, not too hot or too cold, not dangerous, and no toxic chemicals are used in it. For the convenience of reading and checking these packages, they are divided into two categories.

Packs that use physical changes for heat exchange

Most athletes use packs that provide instant cold or heat to treat their injuries (Arzani et al., 2008, p. 33).

The operation of these packages, which are among the first aid tools, is based on the enthalpy of dissolution. These packages consist of a plastic bag inside which there is a small package of water and a round chemical substance. Hitting a plastic bag causes a small packet of water to tear, the chemical substance dissolves in it. If the dissolution of a chemical substance in water is exothermic, the closed temperature will increase and if it is endothermic, the temperature will decrease. Calcium chloride is usually used in hot producing packages and ammonium nitrite is used in cold producing packages. The liquidation process that occurs in each case is:

$$CaCl_2 \xrightarrow{H_2O} Ca^{+}(aq) + 2Cl^{-}(aq) \qquad \Delta H_{dissoln} = -83\ kJ.mol^{-1}$$

Figure 7: Packages producing heat and cold

$$NH_4NO_3 \xrightarrow{H_2O} NH_4^{+}(aq) + NO_3^{-}(aq) \qquad \Delta H_{dissoln} = +26\ kJ.mol^{-1}$$

Experiments show that adding $40g\,CaCl_2\,(0.36mol)$ to $100ml$ of $20^{\circ}C$ water increases the temperature of the water to $90^{\circ}C$. On the other hand, if we add $30g\,NH_4NO_3\,(0.38mol)$ to $100ml$ water $20^{\circ}C$, the water temperature will decrease to $0^{\circ}C$. Hot or cold packs work in about 20 minutes (Rouh Elahi et al., 2009, p. 85).

Packages that use chemical interactions for heat exchange

Exothermic packages sometimes remain warm for 6 hours and can be used again. The way these packages work is that there is a cloth or paper bag with small burners inside a plastic package. This cloth bag contains a mixture of flour, salt, charcoal and sawdust kneaded in water. When an outer plastic layer is removed and a strong shock is applied to the bag and heat is generated. This process is such that iron combines with air oxygen and produces iron oxide. Ordinary iron zinc is formed very slowly and the heat exchanged is not palpable. But in this case, rusting occurs faster due to the presence of moisture, the roundness of iron and the presence of salt (Arzani, 2008, p. 38).

CONCLUSION

From this research, we conclude that in crystalline solids, ions are located in certain places of a network and their movement is limited by vibration around fixed points. Regular crystal structures, whose constituent particles are formed in three dimensions, are called the crystal lattice of that object. The formation of the crystal network is accompanied by the release of energy, and the energy of the network is the amount of energy released during the formation of an ionic solid molecule from its constituent gas ions. The enthalpy of lattice formation of crystalline solids is always positive and its energy can be calculated using the Haber-Bron cycle. In an ionic solid, ions with opposite charges are held together by electrostatic attraction. Ionic solids are mostly soluble in water, and among other liquids, water is the best solvent for ionic and polar compounds. So that the partial positive and negative poles of water establish a relationship with the positive and negative ions of the dissolved substances separate them from the corresponding network and dissolve them in itself. As a result of the dissolution of one substance in another substance, especially water takes place in three stages: separation of solute particles, separation of solvent particles, and mixing of solute and solvent particles. The process of dissolution is associated with energy exchange, which sometimes takes place by absorption and sometimes by release. This energy is obtained by using the enthalpies of the solute substance, solvent, and hydration enthalpy in the form of enthalpy of solution. If the set of mentioned steps gives numbers with a positive sign, the process is endothermic, and if the numbers with a negative sign are obtained, the process is exothermic. Scientists use this property of material dissolution to produce heat-generating and cold-generating packs that are used in medicine and sports.

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